Honors Chemistry Chapter 5 The Periodic Law

This unit covers all of Chapter 5 as well as Chapter 1 Section 3 pgs. 16-20.

Section 1: History of the Periodic Table pgs. 133-137.

Objectives:

1. **Explain the roles of Mendeleev and Moseley in the development of the periodic table.**
2. **Describe the modern periodic table.**
3. **Explain how the periodic law can be used to predict the physical and chemical properties of elements.**
4. **Describe how the elements belonging to a group of the periodic table are inter-related in terms of atomic number.**

Define the following:

1. **periodic law--**

2. **periodic table--**

3. **lanthanide--**

4. **actinide--**

Mendeleev and Chemical Periodicity

In 1869, Russian chemist Dmitri Mendeleev produced the first modern periodic table based on increasing atomic mass. It was a later chemist HGJ Moseley, who modified Mendeleev's work, basing his table on atomic number rather than atomic mass. This led to the modern periodic law, which states that when elements are arranged in order of increasing atomic number their physical and chemical patterns show a periodic (repeating) pattern.

Elements in the periodic table show periodicity or repeating patterns. This periodicity is due to a similar arrangement of electrons around the nucleus.

The Periodic Table

The modern periodic table is an arrangement of the elements in order of their atomic number so that elements with similar properties fall in the same column.

Groups also known as families are arranged in 18 vertical columns.

The horizontal rows are called periods.

There are three different forms of labeling columns in Europe and America, of which we will use two. The older of these two systems uses a Roman numeral with either an A or a B. The IUPAC system, (The International Union of Pure and Applied Chemistry) labels columns with the numerals 1 to 18.
Some groups have additional names.
Group IA is called the alkali metals.
Group IIA is called the alkaline earth metals.
Group VIIA is called the halogens.
Group VIIIA is called the noble gases.
The lanthanides belong to row 6 of the periodic table and include the 14 elements from 58 to 71.
The actinides are part of row 7 of the periodic table and include the elements 90 to 103.
The lanthanides and actinides are set below the periodic table to save space.
Hydrogen is separated from the rest of the elements in the periodic table because of its unique properties. Though hydrogen by its position in the periodic table may appear to be a metal it more closely resembles the nonmetal halogens.

Answer the following questions.

1. **Who discovered the periodic law?**

2. **Who based the periodic table on atomic mass?**

3. **Who based the periodic table on atomic number?**

4. **What period contains atomic numbers 90-103?**

5. **Why are parts of the 6th and 7th period placed below the main section of the periodic table?**

6. **What is another name for a column in the periodic table?**

7. **What is another name for a row in the periodic table?**

Chapter 1 Section 3: Elements pgs. 16-20.
Objectives:
1. **Use a periodic table to name elements, given their symbols.**
2. **Use a periodic table to write the symbols of elements given their name.**
3. **Describe the arrangement of the periodic table.**
4. **List the characteristics that distinguish metals, nonmetals, and metalloids.**

Vocabulary: Define the following.
1. **group--**

2. **family--**

3. **period--**

4. **metal--**

5. **nonmetal--**

6. **metalloid--**

Elements
Elements are pure substances. An element is listed in the periodic table by its chemical symbol.
Students should now mark on their periodic table the chemical names and symbols they will be responsible to memorize in this course.
Metals, Nonmetals, and Semimetals (Metalloids)
Metals--have luster or shine, are good conductors of heat and electricity, are solids at room temperature (exception liquid mercury), are malleable, and ductile. Metals make up most of the elements of earth and are located on the left and center of the periodic table.
Nonmetals--do not have luster, are poor conductors of heat and electricity, are not malleable or ductile, and many are gases at room temperature, though some are sol-

ids, (bromine is an exception and exists as a liquid at room temperature). Nonmetals are located on the right side of the periodic table.
Semimetals (Metalloids)--share properties of metals and nonmetals. There are only a few metalloids, which will be given to you to mark on your periodic table.

Answer the following questions.

1. **Name two metalloids.**
   a.

   b.

2. **Name two nonmetals**
   a.

   b.

3. **The ability of a metal to be drawn into a thin wire is called ______________.**

4. **The ability of a metal to be hammered into thin sheets is called ____________.**

5. **What type of element is considered an insulator?**

6. **What types of purchased home equipment would have metalloids?**

Chapter 5 Section 2: Electron Configuration and the Periodic Law pgs. 138-149.
Objectives:

1. **Describe the relationship between electrons in sublevels and the length of each period of the periodic table.**
2. **Locate and name the four blocks of the periodic table. Explain the reasons for these names.**
3. **Discuss the relationship between group configurations and group numbers.**
4. **Describe the locations in the periodic table and the general properties of the alkali metals, the alkaline earth metals, the halogens, and the noble gases.**

Vocabulary: Define the following.

1. **alkali metals--**

2. alkaline-earth metals--

3. transition elements--

4. main-group elements--

5. halogens--

In general the electron configuration of an atom's highest occupied energy level determines the chemical properties of the atom.

There is organization in both periods and groups.
There are seven periods (rows) of elements. The length of each period is determined by the number of electrons that can occupy the sublevels filled in that period.
See page 138 table 1.
The periodic table can be subdivided into four sublevels blocks, s, p, d, and f.
The period of an element can be determined from the element's electron configuration. Iodine has a configuration of [Kr]$\underline{5}s^2 4d^{10} \underline{5}p^5$, therefore iodine is in the fifth row.

s-block elements Groups1 and 2- these are the reactive metals with a configuration of ns^1 or ns^2.
The s^1 are the alkali metals. The s^2 are the alkaline earth metals.
Both s block elements are so reactive that they are not found as free elements.
Alkaline earth metals are not as reactive as alkali metals. Both elements have lower melting points proceeding down the group.

Hydrogen is a unique element and though it is placed in column one its properties do not fit any group.
Helium is placed in group 18 even though it has 2 electrons because of its stability.

d-block elements Groups 3-12 are also called the transition metals, first appear when n = 3. These elements are of a higher energy than the n level above them, (3d is higher than 4s) so they fill before that level.

The sum of a d block element's s and d electrons is equal to the group number. For example titanium's configuration is [Ar]$3d^2 4s^2$, and has 2 s electrons plus 2 d electrons for a total of 4 electrons and is therefore found in group 4.
d-block metals have all metallic properties but are not as reactive as group 1 and 2 and are often found as free elements.

<u>p-block elements Groups 13-18</u> The p block elements together with the s block elements are called the <u>main-group elements</u>. The number of electrons in the highest occupied p-block element is the group number minus 10.
p-block elements contain all of the nonmetals and metalloids as well as an extremely reactive group called the <u>halogens</u> group 17. <u>Halogens</u> react with metals to form <u>salts.</u>
p-block metals and metalloids are harder than the s-block elements but softer than the d-block elements.
The last group of p-block elements are the <u>noble gases</u>. These elements have complete electron configurations and are not reactive.

f-block elements are found between groups 3 and 4 in the sixth and seventh period. Each row contains 14 elements. These are also called the <u>inner transition metals</u>.
The lanthanides are all metals similar to group 2 metals.
The actinides are all radioactive.

Answer the following.
Given the following electron configurations, a. identify each element b. tell if it is a metal or nonmetal, c. tell if it is possesses high or low reactivity.

	a. name	b. type	c. reactivity
A. [Ar]$4s^1$			
B. [Ne]$3s^2 3p^6$			
C. [He]$2s^2 2p^3$			

2. Write the symbols and atomic numbers for silver, radon, and zinc.

3. Give the group and period number for each element in question 2.
silver

radon

zinc

4. What groups make up the main-group elements?

5. How are all alkali metals stored?

6. Which two groups are the most reactive elements?

7. What is another name for a compound formed between a metal and a halogen?

Section 3: Electron Configuration and Periodic Properties pgs. 150-164.
Objectives:
1. Define atomic and ionic radii, ionization energy, electron affinity, and electronegativity.
2. Compare the periodic trends of atomic radii, ionization energy, and electronegativity, and state the reasons for these variations.
3. Define valence electrons, and state how many are present in atoms of each main-group element.
4. Compare the atomic radii, ionization energies, and electronegativities of the d-block elements with those of the main-group elements.

Vocabulary: Define the following.
1. atomic radius--

2. ion--

3. ionization--

4. ionization energy--

5. electron affinity--

6. cation--

7. anion--

8. valence electrons--

9. electronegativity--

Atomic Radii

The atomic radius is the distance from the center of the atom to its outermost electron or half the distance between identical atoms that are bonded together.

Trends

1. Atoms get larger going down a group.

2. Atoms get smaller moving from left to right.
(This is caused by an increasing number of protons in the nucleus that cause a stronger attraction for the electrons).

Ionization Energy

The ionization energy is the energy needed to remove one of an atom's electrons. High ionization energy means that the atom holds onto its electrons strongly. Low ionization energy means the atom loses electrons easily. When an atom loses or gains an electron it becomes an ion and the process is called ionization.
The ionization energy for an atom is represented by joules/atom and for a large group of atoms it is represented as kilojoules/mole.

Trends

1. Ionization energies decrease as you move down a group.
(Group IA has the lowest ionization energy). This is due to electron shielding that results when the innermost electrons shield the outermost electrons from the full attraction of the positive nucleus.

2. Ionization energies increase as you move across a period.
(The noble gases Group VIIIA have the highest ionization energy).

Note: The trend in ionization energy is opposite the trend in atomic radii.

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Successive ionization energies means that more energy is needed to remove successive electrons from an atom. This is due to less electron-electron repulsion, and a stronger nuclear attraction.

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Teacher's Note on Ions

It was Benjamin Franklin and his experiments with electricity who gave us the terms positive and negative. He later stated he probably should have reversed the names because of the confusion it sometimes caused. In the case of an ion, when an atom gains an electron it becomes negative and when it loses an electron it becomes positive because an electron is a negative particle. Don't confuse the words gain or lose with the fact that ions are opposite the usual meaning of these terms in every-day usage. To gain an electron means negative and to lose an electron means positive.

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Electron Affinity

Electron affinity is the energy change that occurs when an atom gains an electron. Some atoms tend to gain electrons, while others tend to lose electrons. Atoms that want electrons have negative electron affinities, because energy is released. The greater the affinity for electrons is the higher the negative value. Electron affinities are represented by kilojoules/mole.

Trends

1. The nonmetals on the right side of the periodic table have greater negative electron affinities than metals on the left side of the periodic table.

2. Noble gases have the greatest positive electron affinities.

The trend in electron affinity is described by the octet rule, to be discussed momentarily.

Ionic Radii

Trends

1. An atom that loses an electron(s) forms a positive ion called a cation and this resulting ion becomes smaller than the parent atom. (The more electrons lost, the smaller they become).

2. An atom that gains an electron(s) forms a negative ion called an anion and this resulting ion becomes larger than the parent atom. (The more electrons gained, the larger the ion becomes).

The Octet Rule

The octet rule states that atoms tend to gain, lose, (ionic bonds), or share (covalent bonds) electrons in order to acquire a full set of valence electrons; (s and p electrons).

For most atoms the perfect number of valence electrons are 2 electrons in the s orbital and 6 electrons in the p orbitals; resulting in an octet (8).
The nonmetal atoms on the right side of the periodic table tend to gain electrons, becoming negative ions.
The metals on the left side of the periodic table tend to lose electrons becoming positive ions.
The noble gases have the perfect number of electrons (8), and do not form ions or enter into chemical bonds and are therefore for the most part are chemically not reactive.

Electronegativity

Electronegativity is the ability to attract electrons in a chemical bond. Electronegativity is not an amount of energy and therefore does not have any units. Fluorine has the highest electronegativity and cesium and francium have the lowest electronegativity. Fluorine is the most electronegative element and was arbitrarily assigned a value of 4. Values of the other elements are related to this value. Differences between electronegativities will be used in later chapters to determine whether a compound formed between two elements is covalent or ionic.

Trends

1. **Electronegativity increases across a period.**
2. **Electronegativity decrease down a group.**

Answer the following.

1. **What is the most electronegative element?**

2. **What does high ionization mean?**

3. **Why do the atomic radii become smaller moving across a group?**

4. **Why does ionization energy become less moving down a group.**

5. **Which periodic trend deals with the ability to form compounds by subtracting one value for an atom from another atom's value?**

6. **What does a high electron affinity mean?**